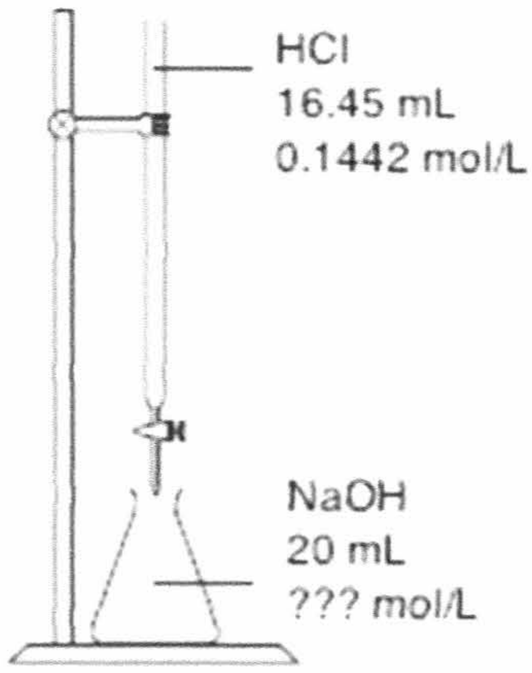
Question 1:

A titration was performed using 20 mL aliquots of NaOH and a standardised solution of 0.1442 mol L^{-1} HCℓ. The NaOH required 16.45 mL of HCℓ to be neutralised. Find the concentration of the NaOH solution.

$$HC\ell + NaOH \rightarrow NaC\ell + H_2O$$



HINT 1: *Sketches like the one to the right can be really useful in making sure you understand what is happening.*

HINT 2: *Always start by working with the 'known' substance. In this case, find the moles of HCℓ.*

$n(HCl) = c \times V$

$= 0.1442 \times 0.01645$

$= 0.002372$ mol

$n(NaOH) = n(HCl)$

$= 0.002372$ mol

$c(NaOH) = n/V$

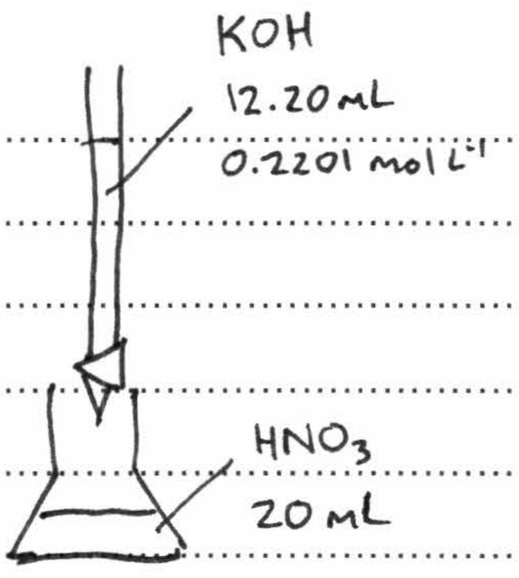
$= 0.002372 / 0.020$

$= 0.1186$ mol L^{-1}

Question 2:

A titration was performed using 20 mL aliquots of nitric acid (HNO_3), with 0.2201 mol L^{-1} KOH in the burette. Three titrations were performed, with an average titre volume of 12.20 mL. Calculate the concentration of the nitric acid.

$$KOH + HNO_3 \rightarrow KNO_3 + H_2O$$



$n(KOH) = c \times V$

$= 0.2201 \times 0.01220$

$= 0.002685$ mol

$n(HNO_3) = n(KOH)$

$= 0.002685$ mol

$c(HNO_3) = n/V$

$= 0.002685 / 0.020$

$= 0.1343$ mol L^{-1}

Question 3:

A solution of NaOH was standardised by reacting with a 0.1008 mol L^{-1} solution of H_2SO_4. An average volume of 18.64 mL of sulfuric acid was required to neutralise the 20.00 mL aliquots of NaOH. Find the concentration of NaOH.

$$H_2SO_4 + 2\ NaOH \rightarrow Na_2SO_4 + 2\ H_2O$$

HINT: *This problem is different from the others because sulfuric acid is a* ***diprotic acid****. There is a 2:1 mole ratio between H_2SO_4 and NaOH. $n(NaOH) = 2 \times n(H_2SO_4)$.*

$n(H_2SO_4) = c \times V$

$= 0.1008 \times 0.01864$

$= 0.001879$ mol

$n(NaOH) = 2 \times n(H_2SO_4)$

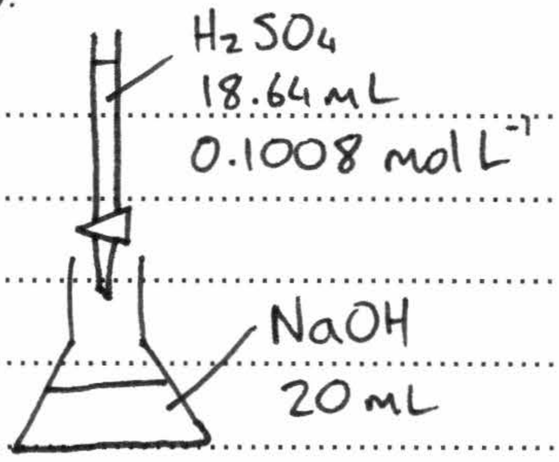
$= 2 \times 0.001879$ mol

$= 0.003758$ mol

$c(NaOH) = n/V$

$= 0.003758/0.020$

$= 0.1879$ mol L^{-1}



Question 4:

A solution of phosphoric acid (H_3PO_4), a triprotic acid, was standardised by reacting with a 0.08555 mol L^{-1} solution of NaOH. An average volume of 27.55 mL of ~~sulfuric~~ phosphoric acid was required to neutralise the 20.00 mL aliquots of NaOH. Find the concentration of phosphoric acid.

$$H_3PO_4 + 3\ NaOH \rightarrow Na_2SO_4 + 2\ H_2O$$

HINT: *Once again, be careful of the mole ratio. It is not 1:1.*

$n(NaOH) = c \times V$

$= 0.08555 \times 0.020$

$= 0.001711$ mol

$n(H_3PO_4) = \frac{1}{3} \times n(NaOH)$

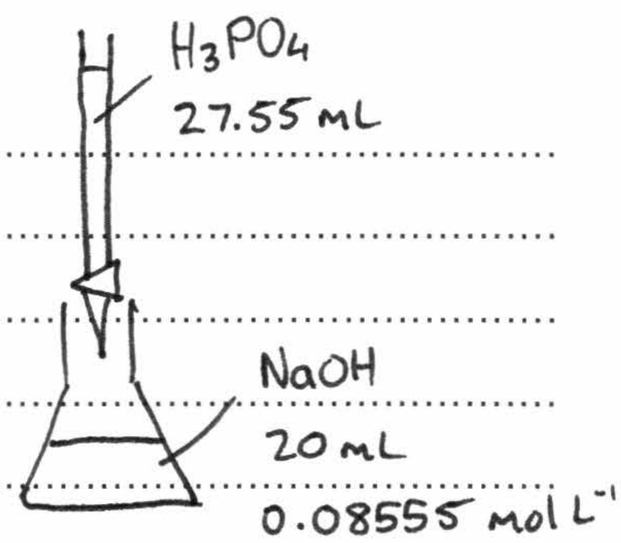
$= \frac{1}{3} \times 0.001711$

$= 0.000570$ mol

$c(H_3PO_4) = n/V$

$= 0.000570/0.02755$

$= 0.02070$ mol L^{-1}

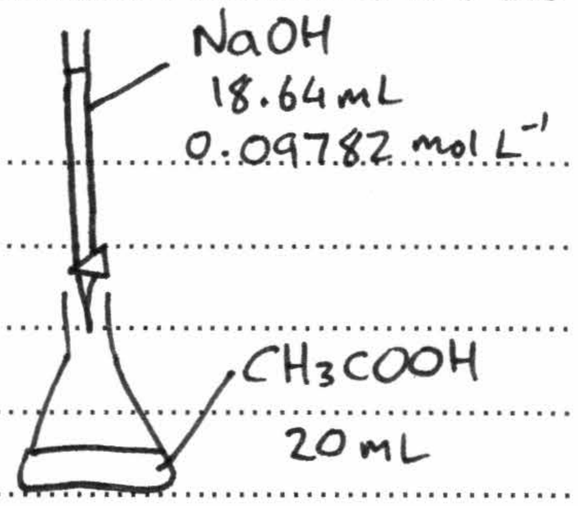


Question 5:

The main ingredient in vinegar is acetic acid (CH_3COOH), a monoprotic acid. A solution of acetic acid was standardised by reacting with a 0.09782 mol L^{-1} solution of NaOH. An average volume of 18.64 mL of NaOH was required to neutralise the 20.00 mL aliquots of CH_3COOH. Find the concentration of acetic acid.

HINT: *No equation was written for this reaction. You should write a balanced reaction to find the mole ratio.*

$NaOH + CH_3COOH \rightarrow NaCH_3COO + H_2O$.



$n(NaOH) = c \times V$

$= 0.09782 \times 0.01864$

$= 0.001823$ mol.

$n(CH_3COOH) = n(NaOH)$

$= 0.001823$ mol

$c(CH_3COOH) = n/V$

$= 0.001823 / 0.020$

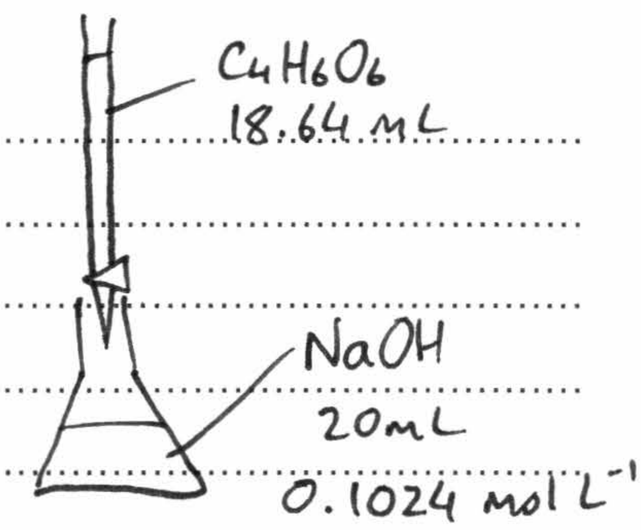
$= 0.09117$ mol L^{-1}

Question 6:

The main ingredient in wine is tartaric acid ($C_4H_6O_6$), a diprotic acid. A solution of tartaric acid was standardised by reacting with a 0.1024 mol L^{-1} solution of NaOH. An average volume of 18.64 mL of tartaric acid was required to neutralise the 20.00 mL aliquots of NaOH. Find the concentration of tartaric acid.

loses 2 H^+
∴ reacts with 2 OH^-

$C_4H_6O_6 + 2NaOH \rightarrow C_4H_4O_6^{2-} + 2Na^+ + 2H_2O$



$n(NaOH) = c \times V$

$= 0.1024 \times 0.020$

$= 0.002048$ mol

$n(C_4H_6O_6) = \frac{1}{2} \times n(NaOH)$

$= 0.001024$ mol

$c(C_4H_6O_6) = n/V$

$= 0.001024 / 0.01864$

$= 0.05494$ mol L^{-1}

Question 7:

A titration was performed using 20 mL aliquots of 0.04056 mol L^{-1} solution of sodium carbonate. The sodium carbonate was titrated against a solution of nitric acid with a concentration of approximately 0.1 mol L^{-1}. The following information was recorded for the titration.

	Volume of HNO_3 solution required			
	1	2	3	4
Initial volume (mL)	0.45	1.20	0.05	12.40
Final volume (mL)	20.80	20.40	19.25	31.65
Titre volume (mL)	~~20.35~~ Rough / Outlier	19.20	19.20	19.25

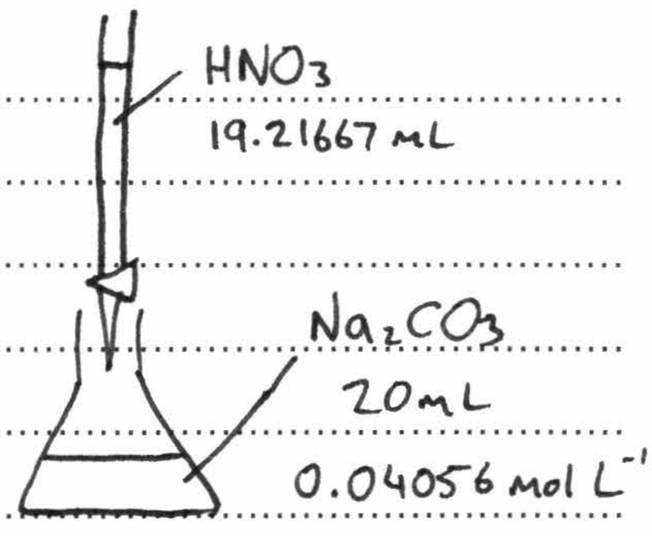
Calculate the concentration of the nitric acid solution.

Hint: *You will need to calculate the <u>average titre volume</u> to use in your calculation. Don't include any outliers in your average. Only include concordant (consistent) values. i.e. within ± 0.1 mL.*

$Na_2CO_3 + 2\ HNO_3 \rightarrow 2\ NaNO_3 + H_2O + CO_2$

$V(HNO_3) = \frac{19.20 + 19.20 + 19.25}{3}$

$= 19.21667$ mL



$n(Na_2CO_3) = c \times V$

$= 0.04056 \times 0.020$

$= 0.000811$ mol

$n(HNO_3) = 2 \times n(Na_2CO_3)$

$= 2 \times 0.000811$

$= 0.001622$ mol

$c(HNO_3) = n / V$

$= 0.001622 / 0.01921667$

$= 0.08443$ mol L^{-1}

Question 8:

A solution of sodium hydroxide was standardised by titrating against 1.067 mol L^{-1} H_2SO_4. The titration was performed using 20 mL aliquots of H_2SO_4 with the NaOH in the burette. The following information was recorded for the titration:

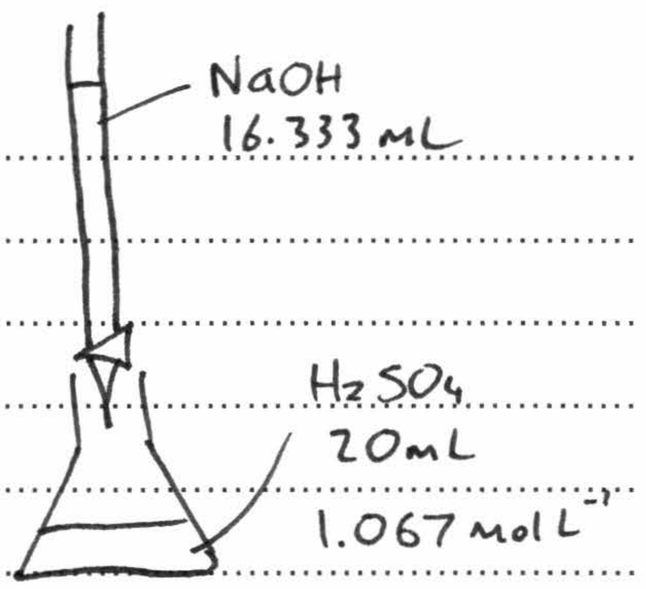
	Volume of NaOH solution required				
	1	2	3	4	5
Initial volume (mL)	0.20	1.65	1.60	17.20	12.40
Final volume (mL)	17.85	18.05	17.90	36.15	28.70
Titre volume (mL)	17.65	16.40	16.30	18.95	16.30

Calculate the concentration of the sodium hydroxide solution.

$H_2SO_4 + 2\,NaOH \rightarrow Na_2SO_4 + H_2O$

Avg titre = $\frac{16.40 + 16.30 + 16.30}{3}$

= 16.333 mL



$n(H_2SO_4) = c \times V$

$= 1.067 \times 0.020$

$= 0.02134$ mol

$n(NaOH) = 2 \times n(H_2SO_4)$

$= 2 \times 0.02134$

$= 0.04268$ mol

$c(NaOH) = n/V$

$= 0.04268 / 0.016333$

$= 2.613$ mol L^{-1}

Question 9:

A solution of household ammonia was analysed by titration to determine the concentration of ammonia in the solution. Prior to analysis, the ammonia was **diluted** by pipetting a 10.00 mL aliquot into a 500 mL volumetric flask and making up to the mark using distilled water. The titration was then performed using 20.00 mL aliquots of the **diluted** ammonia solution in the conical flask, and 0.07844 mol L^{-1} hydrochloric acid in the burette.

The following information was recorded for the titration:

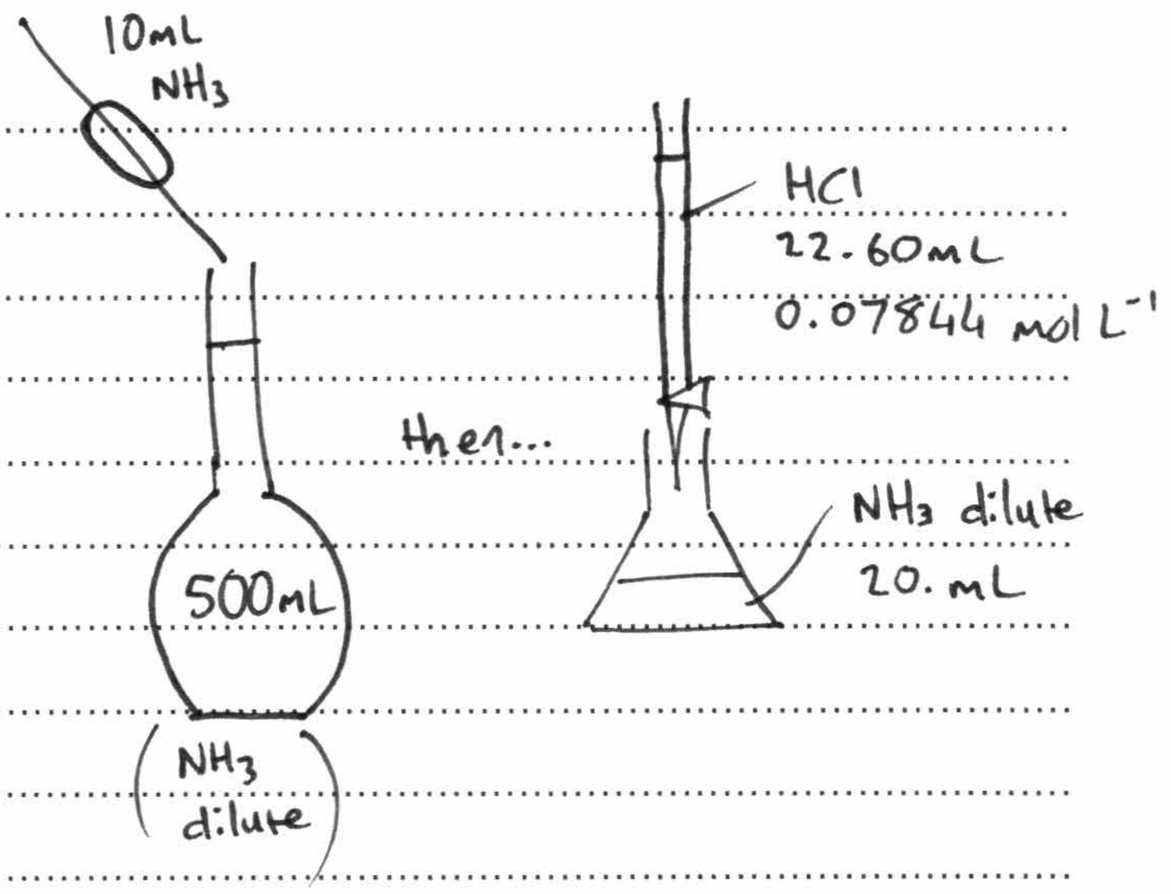
	Volume of HCℓ solution required			
	1	2	3	4
Initial volume (mL)	0.75	7.20	4.30	4.95
Final volume (mL)	23.40	29.80	26.30	27.50
Titre volume (mL)	22.65	22.60	22.00	22.55

Calculate the concentration (in mol L^{-1}) of ammonia in the **original (undiluted)** solution.

HINT: *First, calculate the concentration of diluted ~~vinegar~~ ammonia using information about the titration. Then you can use the formula $c_1 \times V_1 = c_2 \times V_2$ to calculate the concentration of the undiluted ammonia solution.*

$HCl + NH_3 \rightarrow NH_4Cl$

Avg titre = 22.60 mL.



$n(HCl) = c \times V$

$= 0.07844 \times 0.02260$

$= 0.001773$ mol.

$n(NH_3 \text{ dilute}) = n(HCl)$

$= 0.001733$ mol.

$c(NH_3 \text{ dilute}) = n/V$

$= 0.001733 / 0.020$

$= 0.088637$ mol L^{-1}

$c(NH_3 \text{ original}) = c(NH_3 \text{ dilute}) \times \frac{V(\text{dilute})}{V(\text{original})}$

$= 0.088637 \times \frac{500 \text{ mL}}{10 \text{ mL}}$

$= 4.432$ mol L^{-1}

Question 10:

A solution of vinegar was analysed by titration to determine the concentration of acetic acid. Prior to analysis, the vinegar was diluted by pipetting a 25.00 mL aliquot into a 250 mL volumetric flask and making up to the mark using distilled water. The titration was then performed using 20.00 mL aliquots of the 0.1005 mol L^{-1} NaOH in the conical flask, and diluted vinegar in the burette.

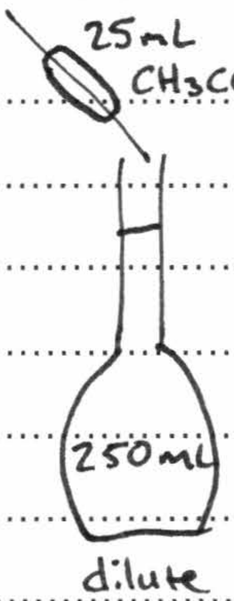
The following information was recorded for the titration:

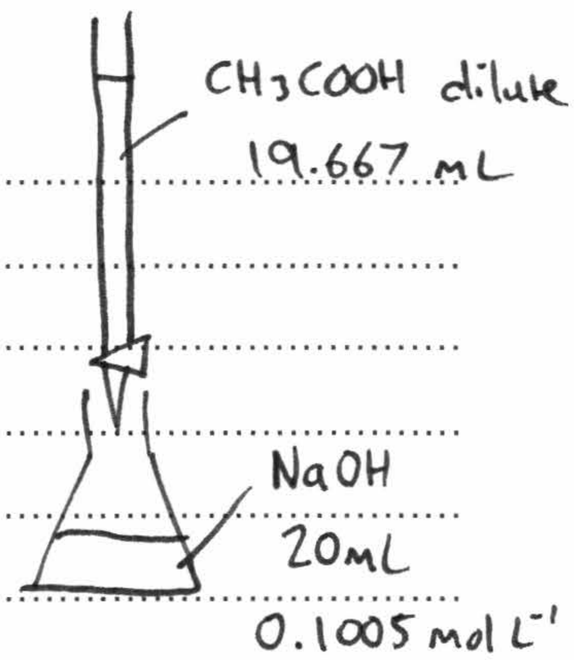
	Volume of diluted CH_3COOH solution required			
	1	2	3	4
Initial volume (mL)	0.25	4.55	12.60	6.60
Final volume (mL)	20.30	24.25	32.30	26.20
Titre volume (mL)	20.05	19.70	19.70	19.60

×

Calculate the concentration (in mol L^{-1}) of acetic acid in the vinegar.

$CH_3COOH + NaOH \rightarrow NaCH_3COO + H_2O$





Avg titre = 19.667 mL

$n(NaOH) = c \times V$

$= 0.1005 \times 0.020$

$= 0.00201$ mol.

$n(CH_3COOH\ dilute) = n(NaOH)$

$= 0.00201$ mol.

$c(CH_3COOH\ dilute) = n/V$

$= 0.00201/0.019667$

$= 0.102203$ mol L^{-1}

$c(CH_3COOH\ original) = c(CH_3COOH\ dilute) \times \frac{250\ mL}{25\ mL}$

$= 0.102203 \times 250/25$

$= 1.022$ mol L^{-1}

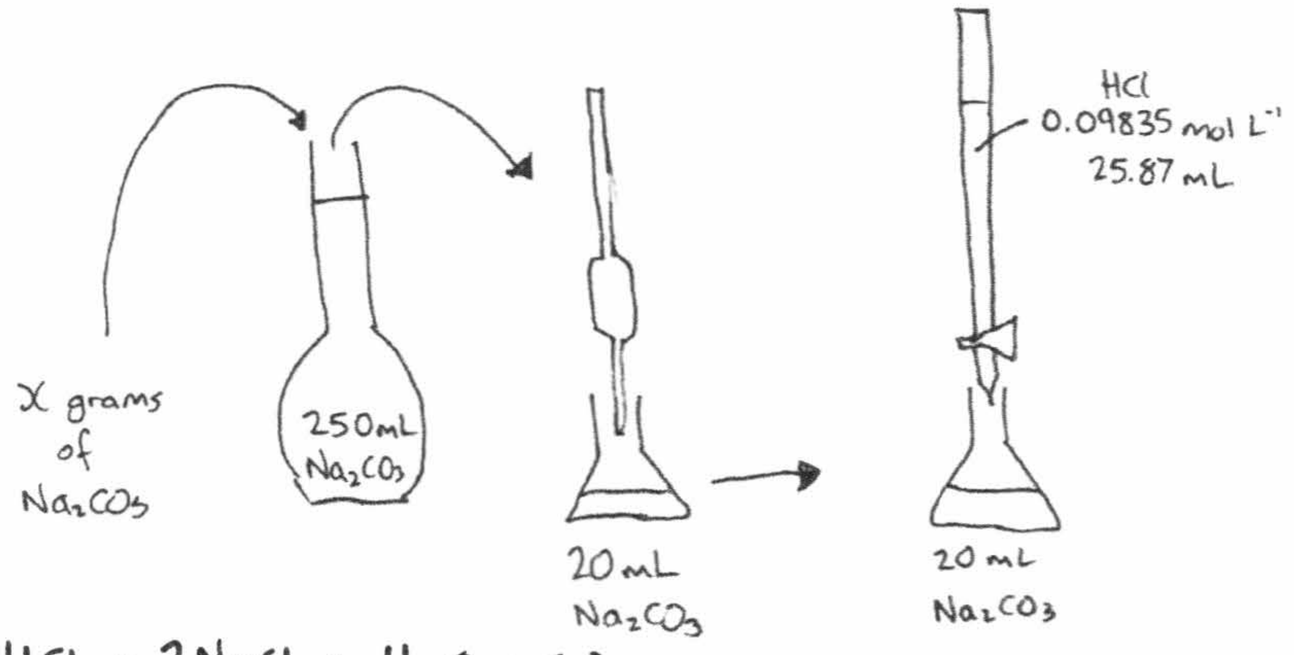
Question 11:

A solution of Na_2CO_3 was prepared by dissolving an unknown mass of anhydrous Na_2CO_3 into 250.0 mL of water. 20.00 mL aliquots of the resulting solution were titrated against 0.09835 mol L^{-1} HCℓ. The average titre volume was 25.87 mL.

Calculate the mass of Na_2CO_3 dissolved in the volumetric flask.

HINT: *Use your normal titration calculations to find the number of moles of Na_2CO_3 in the conical flask, and then from this find the number of moles of Na_2CO_3 in the volumetric flask.*

Remember, a simple diagram of the steps can aid in understanding what is happening.



$Na_2CO_3 + 2HCl \rightarrow 2NaCl + H_2O + CO_2$

$n(HCl) = c \times V$

$= 0.09835 \times 0.02587$

$= 0.002544$ mol.

$n(Na_2CO_3) = \frac{1}{2} \times n(HCl)$

$= \frac{1}{2} \times 0.002544$

$= 0.001272$ mol. (in 20 mL aliquot)

$n(Na_2CO_3 \text{ in vol. flask}) = n(Na_2CO_3 \text{ in conical}) \times \frac{250}{20}$

$= 0.015902$ mol.

$M(Na_2CO_3) = n \times M$

$= 0.015902 \times 105.99$

$= 1.685$ g.

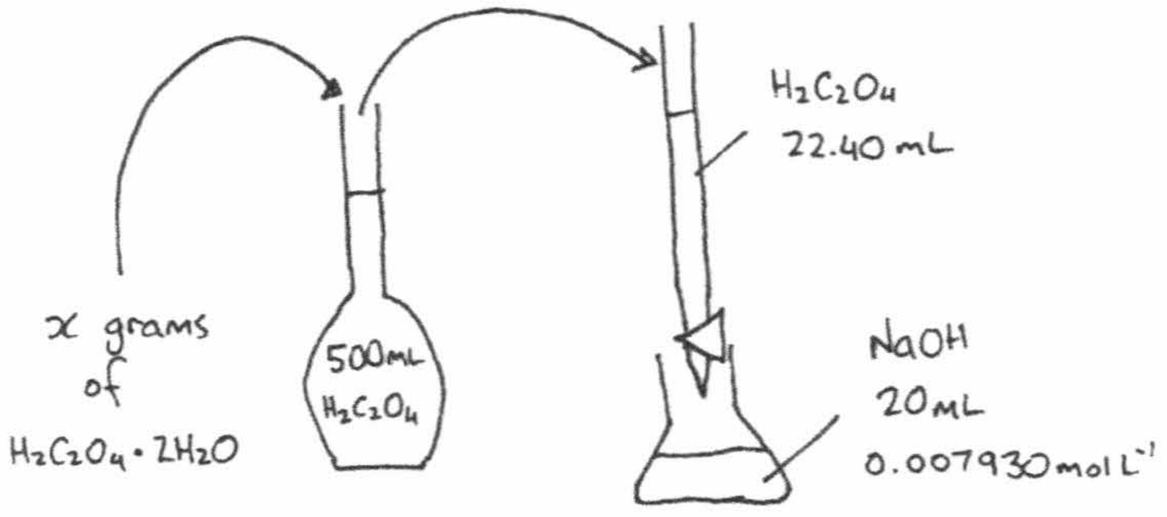
Question 12:

A solution of oxalic acid ($H_2C_2O_4$) was prepared by dissolving an unknown mass of oxalic acid dihydate ($H_2C_2O_4.2H_2O$) in a 500 mL volumetric flask. The oxalic acid was placed in the burette, and was titrated against 20.00 mL aliquots of 0.007930 mol L^{-1} NaOH solution. The average titre volume was ~~25.87~~ 22.40 mL.

Calculate the mass of oxalic acid dihydrate dissolved in the 500 mL volumetric flask.

Note: Oxalic acid ($H_2C_2O_4$) is a diprotic acid.

HINT: *Use your normal titration calculations to find the number of moles of $H_2C_2O_4$ in the 22.40 mL titre volume, and from this find the number of moles of $H_2C_2O_4$ in the volumetric flask.*



$2NaOH + H_2C_2O_4 \rightarrow Na_2C_2O_4 + 2H_2O$

(Sodium oxalate)

$n(NaOH) = c \times V$

$= 0.007930 \times 0.020$

$= 0.000159$ mol.

$n(H_2C_2O_4) = \frac{1}{2} \times n(NaOH)$

$= 0.0000793$ mol. (in 22.40 mL titre)

$n(H_2C_2O_4 \text{ in vol. flask}) = n(H_2C_2O_4 \text{ titre}) \times \frac{500}{22.40}$

$= 0.0017701$ mol.

$n(H_2C_2O_4 \cdot 2H_2O) = n(H_2C_2O_4 \text{ in vol. flask})$

$= 0.0017701$ mol.

$m(H_2C_2O_4 \cdot 2H_2O) = n \times M$

$= 0.0017701 \times 126.068$

$= 0.2232$ g.

Question 13:

Washing soda is an impure mixture which mainly contains sodium carbonate decahydrate ($Na_2CO_3.10H_2O$).

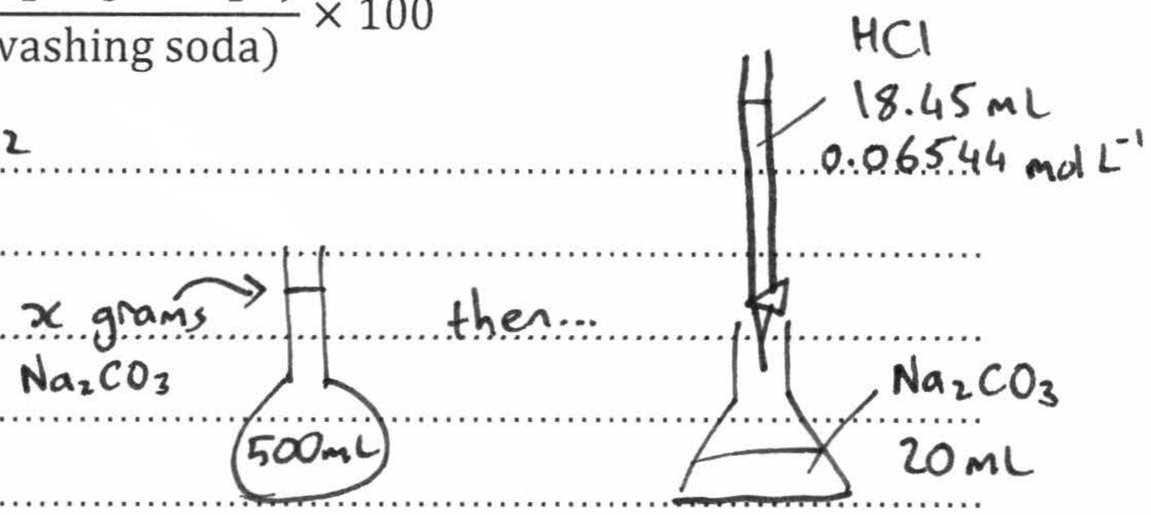
4.81 g of washing soda was dissolved in 500 mL of water and the flask was inverted multiple times to ensure thorough mixing. 20.00 mL aliquots of this solution were then titrated against 0.06544 mol L^{-1} hydrochloric acid using methyl orange indicator. An average titre volume of 18.45 mL was required to reach the end point.

Calculate the percentage by mass of sodium carbonate decahydrate in the washing soda sample.

HINT: *You will need to find the mass of $Na_2CO_3.10H_2O$ in the sample using information from the titration, and then use the following formula:*

$$\%Na_2CO_3.10H_2O = \frac{m(Na_2CO_3.10H_2O)}{m(\text{washing soda})} \times 100$$

$Na_2CO_3 + 2HCl \rightarrow 2NaCl + H_2O + CO_2$



$n(HCl) = c \times V$

$= 0.06544 \times 0.01845$

$= 0.0012074$ mol.

$n(Na_2CO_3) = 1/2 \times n(HCl)$

$= 0.0006037$ mol (in 20 mL aliquot)

$n(Na_2CO_3 \text{ in vol. flask}) = n(Na_2CO_3 \text{ in aliquot}) \times \frac{500}{20}$

$= 0.0150921$ mol.

$n(Na_2CO_3 \cdot 10H_2O) = n(Na_2CO_3 \text{ in vol. flask})$

$= 0.0150921$ mol

$m(Na_2CO_3 \cdot 10H_2O) = n \times M$

$= 0.015091 \times 286.15$

$= 4.3186$ g

$\%Na_2CO_3 \cdot 10H_2O = \frac{m(Na_2CO_3 \cdot 10H_2O)}{m(\text{Washing soda})} \times 100$

$= \frac{4.3186}{4.81} \times 100$

$= 89.8\%$.

Question 14:

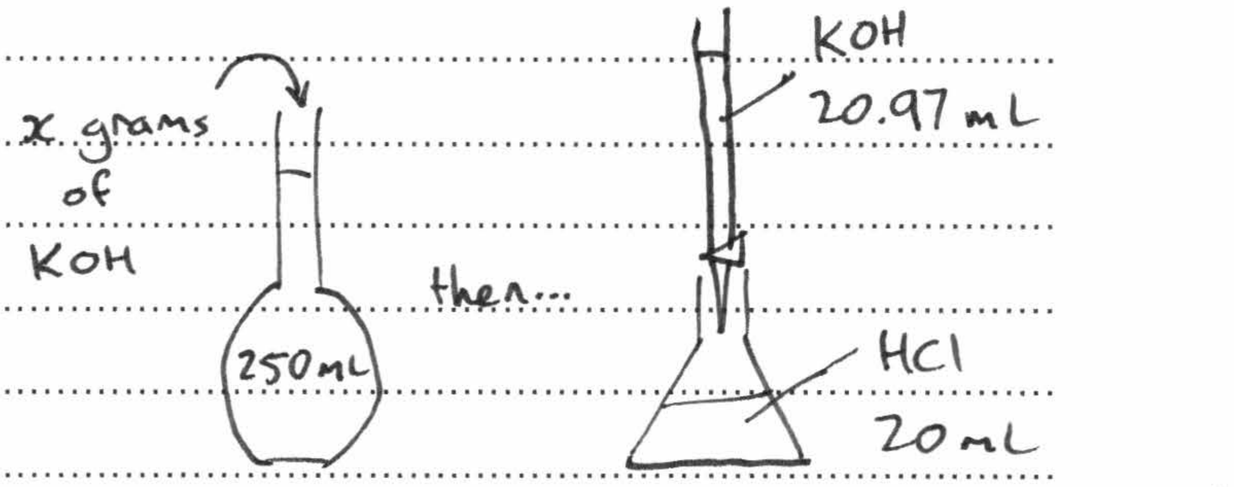
Potassium hydroxide (KOH) cannot be used as a primary standard because it is deliquescent (absorbs water moisture from the air).

A solution of potassium hydroxide was prepared by dissolving ~~1.79~~ 1.32 g of potassium hydroxide pellets in a 250 mL volumetric flask. A small amount of the solution was used to rinse a burette, and then the burette was filled with the potassium hydroxide solution.

Separately, 20.00 mL aliquots of 0.09667 mol L^{-1} HCℓ were pipetted into clean conical flasks, and 2-3 drops of bromothymol blue indicator was added. KOH solution was added until the bromothymol blue changed colour from yellow to green. On average, 20.97 mL of KOH solution was required to reach the end point.

Calculate the %KOH and $\%H_2O$ in the potassium hydroxide pellets used to make the KOH solution.

$HCl + KOH \rightarrow KCl + H_2O$



$n(HCl) = c \times V$

$= 0.09667 \times 0.020$

$= 0.0019334$ mol.

$n(KOH) = n(HCl)$

$= 0.0019334$ mol (in 20.97 mL titre)

$n(KOH \text{ in vol. flask}) = n(KOH \text{ in titre}) \times \frac{250}{20.97}$

$= 0.0230496$ mol.

$m(KOH) = n \times M$

$= 0.0230496 \times 56.108$

$= 1.293$g

$\% KOH = \frac{m(KOH)}{m(KOH \text{ pellets})} \times 100$

$= \frac{1.293}{1.32} \times 100$

$= 98.0\%$

$\therefore \% H_2O = 2.0\%$

Question 15:

The acidity of wine is due mainly to potassium tartrate (cream of tartar), a weak monoprotic acid with a molar mass of 188 g mol^{-1}. Three 50 mL samples of wine were titrated with 0.012 mol L^{-1} NaOH. The results of the titrations are shown below:

	Volume of NaOH solution required			
	1	2	3	4
Initial volume (mL)	0.0	0.0	0.0	0.1
Final volume (mL)	12.65	9.70	10.10	10.20
Titre volume (mL)	12.65	9.70	10.10	10.10

a) Calculate the moles of potassium tartrate in 50 mL of wine
b) Calculate the concentration of potassium tartrate in g L^{-1}.

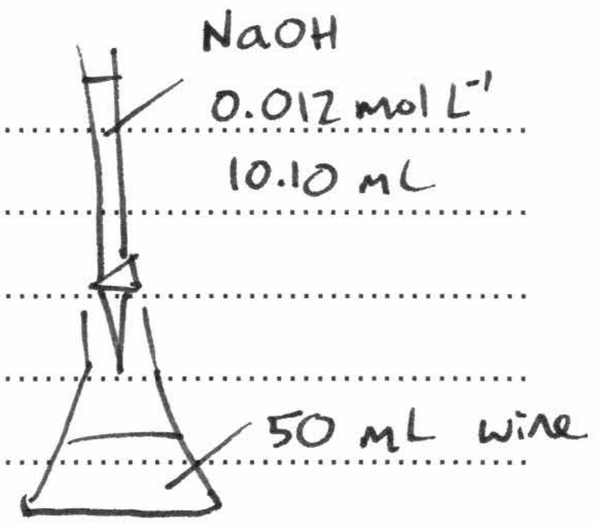
HINT: *Even though you don't have the chemical formula for potassium tartrate, the question stated that it is monoprotic and provided the molar mass, which is all that you need to solve the question. To answer part b, first find the mass of potassium tartrate in 50 mL of the wine, and then divide this by the volume (in litres).*

a) Avg titre = 10.10 mL

$n(NaOH) = c \times V$

$= 0.012 \times 0.01010$

$= 0.0001212$ mol.



Potassium tartrate is monoprotic.

$\therefore n(\text{potassium tartrate}) = n(NaOH)$

$= 0.0001212$ mol. (in 50 mL of wine)

b) $m(\text{potassium tartrate}) = n \times M$

$= 0.0001212 \times 188$

$= 0.022786$ g.

$\text{concentration of potassium tartrate} = \frac{0.022786 \text{ g}}{0.050 \text{ L}}$

$= 0.456 \text{ g L}^{-1}$

Question 16:

A solution of household vinegar was analysed to determine the acetic acid content. 10.00 mL of vinegar was added to a 250 mL volumetric flask and distilled water was added until the total volume of the solution was 250 mL. 20.00 mL aliquots of the diluted vinegar solution were pipetted into clean conical flasks, and 2-3 drops of phenolphthalein was added.

On average, 16.70 mL of 0.0500 mol L^{-1} NaOH was required to neutralise the 20.00 mL aliquots of diluted vinegar.

a) Calculate the mass of CH_3COOH in 10 mL of the original undiluted vinegar
b) Given that the density of household vinegar is 1.05 g/mL, calculate the percentage by mass of acetic acid in the household vinegar.

HINT: *You should be able to solve part a using skills you have developed through other questions. For part b, you are going to need to use the formula:*

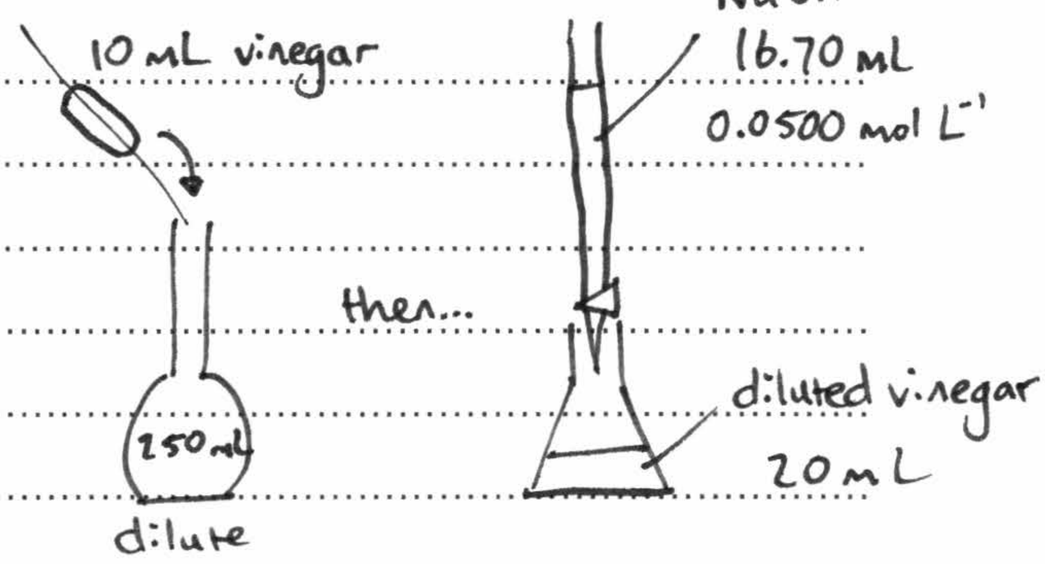
$$\%CH_3COOH = \frac{m(CH_3COOH)}{m(\text{vinegar})} \times 100$$

You will need to use the density to figure out the mass of 10 mL of undiluted vinegar. The equation for density is…

$$\text{density} = \frac{\text{mass}}{\text{volume}}$$

By rearranging this equation you should be able to calculate the mass of 10 mL of vinegar.

a) $CH_3COOH + NaOH \rightarrow NaCH_3COO + H_2O$.



$n(NaOH) = c \times V$

$= 0.0500 \times 0.01670$

$= 0.000835$ mol.

$n(CH_3COOH) = n(NaOH)$

$= 0.000835$ mol (in 20 mL dilute)

$n(CH_3COOH \text{ in vol. flask}) = n(CH_3COOH \text{ in conical}) \times {}^{250}/_{20}$

$= 0.000835 \times {}^{250}/_{20}$

$= 0.0104375$ mol.

$n(CH_3COOH \text{ in 10 mL original}) = n(CH_3COOH \text{ in vol. flask})$

$= 0.0104375$ mol.

b) $m(CH_3COOH) = n \times M$

$= 0.0104375 \times 60.052$

$= 0.62679$ g.

$m(\text{vinegar}) = \text{density} \times \text{volume}$

$= 1.05 \text{ g/mL} \times 10 \text{ mL}$

$= 10.5$ g.

$\%CH_3COOH = \frac{m(CH_3COOH)}{m(\text{vinegar})} \times 100$

$= \frac{0.62679}{10.5} \times 100$

$= 5.97\%$

Question 16:

A chemist analysed aspirin tablets for quality control. The initial step of the analysis was the standardisation of a NaOH solution. Three 25.00 mL samples of a 0.1034 mol L^{-1} solution of standardised HCℓ were titrated with the NaOH solution. The average volume required for neutralisation was 25.75 mL.

Three conical flasks were prepared each containing a mixture of 25 mL of water and 10 mL of ethanol. One aspirin tablet was dissolved in each flask. The aspirin in each solution was then titrated with the standardised NaOH solution according to the following equation:

$$C_9H_8O_4(aq) + NaOH(aq) \rightarrow C_9H_7O_4Na(aq) + H_2O(\ell)$$

The following titration results were obtained.

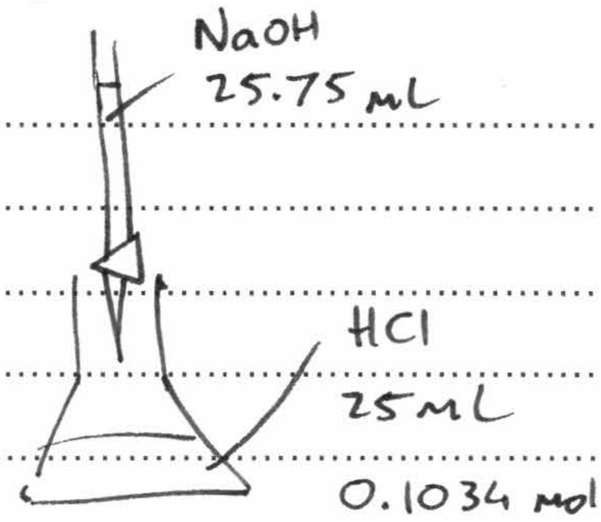
Tablet	Volume of NaOH (mL)
1	16.60
2	16.50
3	16.55

Avg titre = 16.55 mL.

Calculate the average mass (in mg) of aspirin in each tablet.

HINT: *There are two titrations here. Start by finding the c(NaOH) using the HCℓ titration.*

Titration #1: $NaOH + HCl \rightarrow NaCl + H_2O$



$n(HCl) = c \times V$

$= 0.1034 \times 0.025$

$= 0.002585$ mol.

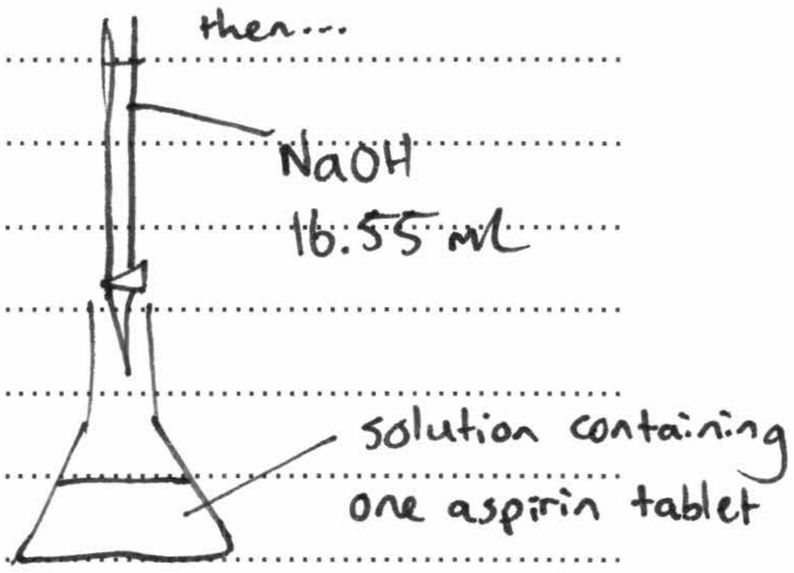
$n(NaOH) = n(HCl)$

$= 0.002585$ mol.

$c(NaOH) = n/V$

$= 0.002585/0.02575$

$= 0.10039$ mol L^{-1}



Titration #2:

$n(NaOH) = c \times V$

$= 0.10039 \times 0.01655$

$= 0.0016614$ mol.

$n(C_9H_8O_4) = n(NaOH)$

$= 0.0016614$ mol

$m(C_9H_8O_4) = n \times M$

$= 0.0016614 \times 180.154$

$= 0.299$ g

$= 299$ mg

Question 17:

A student was asked to determine the mass, in grams, of calcium carbonate present in a 0.125 g sample of chalk. The student placed the chalk in a 250 mL conical flask and added 50.00 mL of 0.200 mol L^{-1} HCℓ using a volumetric pipette. The solution was stirred until no further reaction occurred.

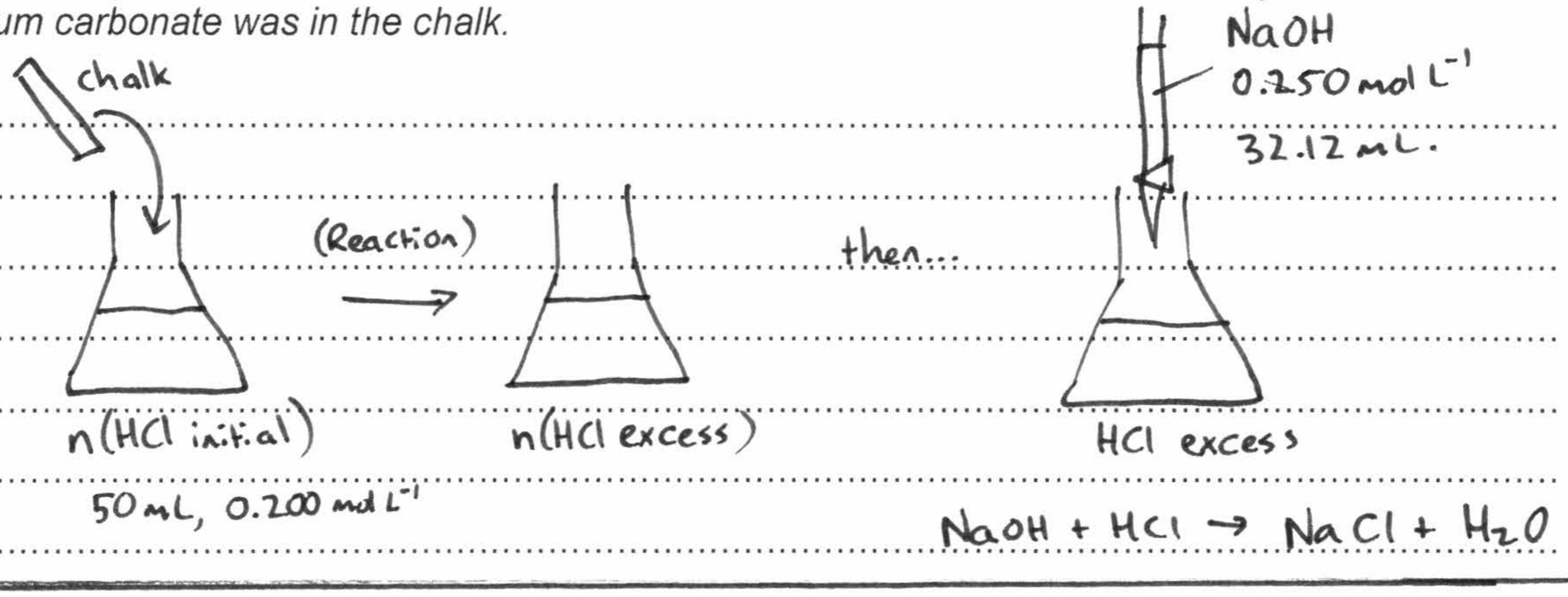
The student then titrated the excess HCℓ against 0.250 mol L^{-1} NaOH. The average NaOH titre volume was 32.12 mL.

Calculate the mass of calcium carbonate, in grams, present in the chalk sample.

HINT: *This strategy is called a 'back titration'. From the first paragraph you can calculate the n(HCℓ) that were initially added to the flask. Some of this HCℓ would have reacted with the calcium carbonate, but some was left over. The n(HCℓ) left over was measured in the NaOH titration.*

n(HCℓ reacted) = n(HCℓ initial) – n(HCℓ excess)

Find out how many moles of HCℓ reacted with the calcium carbonate, and use that to figure out how much calcium carbonate was in the chalk.



$n(HCl\ initial) = c \times V$
$= 0.200 \times 0.050$
$= 0.0100$ mol.

$n(NaOH) = c \times V$
$= 0.250 \times 0.03212$
$= 0.00803$ mol.

$n(HCl\ excess) = n(NaOH)$
$= 0.00803$ mol.

$n(HCl\ reacted) = n(HCl\ initial) - n(HCl\ excess)$
$= 0.0100 - 0.00803$ mol
$= 0.00197$ mol.

$CaCO_3 + 2HCl \rightarrow CaCl_2 + CO_2 + H_2O$.

$n(CaCO_3) = \frac{1}{2} \times n(HCl\ reacted)$
$= \frac{1}{2} \times 0.00197$
$= 0.000985$ mol.

$m(CaCO_3) = n \times M$
$= 0.000985 \times 100.09$
$= 0.0986$ g.

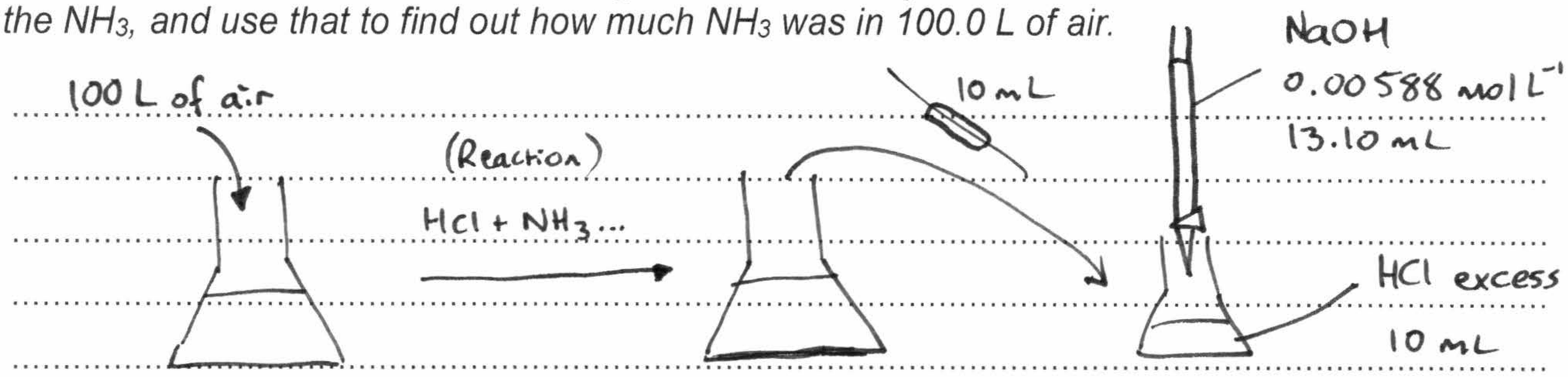
Question 18:

US federal regulations set the upper limit of ammonia (NH_3) in air in a work environment of 38.5 µg (38.5×10^{-6} g) per litre of air. To test the air quality in a manufacturing plant, 100.0 L of air was bubbled through 100 mL of 0.0105 mol L^{-1} HCℓ.

10.0 mL aliquots of the reacted mixture is then titrated with 0.00588 mol L^{-1} NaOH. On average, 13.10 mL of NaOH was required to neutralise the excess HCℓ.

Calculate the concentration of ammonia in the air sample (in µg/L) and state whether or not the manufacturer is in compliance with regulations.

HINT: *This is another back titration. Figure out how many moles of HCℓ must have reacted with the NH_3, and use that to find out how much NH_3 was in 100.0 L of air.*



$n(NaOH) = c \times V$

$= 0.00588 \times 0.0130$

$= 7.644 \times 10^{-5}$ mol.

$n(HCl\ excess) = n(NaOH)$

$= 7.644 \times 10^{-5}$ mol (in 10 mL)

$n(HCl\ excess\ in\ 100\ mL) = n(HCl\ in\ 10\ mL) \times \frac{100}{10}$

$= 7.644 \times 10^{-4}$ mol.

$n(HCl\ initial\ in\ 100\ mL) = c \times V$

$= 0.0105 \times 0.100$

$= 0.00105$ mol.

$n(HCl\ reacted) = n(HCl\ initial) - n(HCl\ excess)$

$= 0.00105 - 7.644 \times 10^{-4}$

$= 2.856 \times 10^{-4}$ mol.

$HCl + NH_3 \rightarrow NH_4Cl$.

$n(NH_3) = n(HCl\ reacted)$

$= 2.856 \times 10^{-4}$ mol.

$m(NH_3) = n \times M$

$= 2.856 \times 10^{-4} \times 17.034$

$= 0.004865$ g

$= 4865$ µg.

$concentration = \frac{4865\ \mu g}{100\ L}$

$= 48.65$ µg/L

This is above the max. limit of 38.5 µg/L.

∴ manufacturer is not in compliance.